



AP[®] Chemistry 2002 Sample Student Responses Form B

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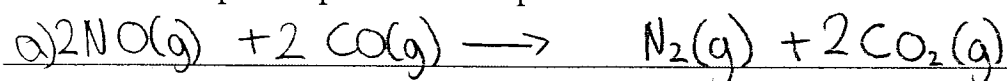
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3. Nitrogen monoxide, $\text{NO}(g)$, and carbon monoxide, $\text{CO}(g)$, are air pollutants generated by automobiles. It has been proposed that under suitable conditions these two gases could react to form $\text{N}_2(g)$ and $\text{CO}_2(g)$, which are components of unpolluted air.

- (a) Write a balanced equation for the reaction described above. Indicate whether the carbon in CO is oxidized or whether it is reduced in the reaction. Justify your answer.
- (b) Write the expression for the equilibrium constant, K_p , for the reaction.
- (c) Consider the following thermodynamic data.

	<u>NO</u>	<u>CO</u>	<u>CO₂</u>
ΔG_f° (kJ mol ⁻¹)	+86.55	-137.15	-394.36

- (i) Calculate the value of ΔG° for the reaction at 298 K.
- (ii) Given that ΔH° for the reaction at 298 K is -746 kJ per mole of $\text{N}_2(g)$ formed, calculate the value of ΔS° for the reaction at 298 K. Include units with your answer.
- (d) For the reaction at 298 K, the value of K_p is 3.3×10^{120} . In an urban area, typical pressures of the gases in the reaction are $P_{\text{NO}} = 5.0 \times 10^{-7}$ atm, $P_{\text{CO}} = 5.0 \times 10^{-5}$ atm, $P_{\text{N}_2} = 0.781$ atm, and $P_{\text{CO}_2} = 3.1 \times 10^{-4}$ atm.
- (i) Calculate the value of ΔG for the reaction at 298 K when the gases are at the partial pressures given above.
- (ii) In which direction (to the right or to the left) will the reaction be spontaneous at 298 K with these partial pressures? Explain.



$$b) K_p = \frac{[P_{\text{CO}_2}]^2 [P_{\text{N}_2}]}{[P_{\text{CO}}]^2 [P_{\text{NO}}]^2}$$

$$\begin{aligned} c) i) \Delta G^\circ &= n(\Delta G_f^\circ)_{\text{products}} - n(\Delta G_f^\circ)_{\text{reactants}} \\ &= (2 \times -394.36) - (2 \times -137.15 + 2 \times 86.55) \\ &= -687.52 \text{ kJ} \end{aligned}$$

$$ii) \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \Rightarrow \Delta S^\circ = \frac{\Delta H^\circ - \Delta G^\circ}{T} = \frac{-746 \text{ kJ} + 687.52 \text{ kJ}}{298 \text{ K}} = \frac{-58.48 \text{ kJ}}{298 \text{ K}}$$

$$= -0.196 \text{ kJ/K} = -196 \text{ J/K}$$

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$$d) i) \Delta G = \Delta G^\circ + RT \ln K_p$$

$$K_p = \frac{0.781(3.1 \times 10^{-4})^2}{(5 \times 10^{-5})^2(5 \times 10^{-7})^2} = 1.20 \times 10^{14}$$

$$= -687.52 \text{ kJ} + (8.31 \times 10^{-3})(298) \ln 1.20 \times 10^{14}$$

$$= -607.24 \text{ kJ}$$

ii) To the left, because ΔG is negative at 298K; this means that reaction will favour moving to the left. The reaction is considered spontaneous if ΔG is -ve. Spontaneous if it produces the products.

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3. Nitrogen monoxide, $\text{NO}(g)$, and carbon monoxide, $\text{CO}(g)$, are air pollutants generated by automobiles. It has been proposed that under suitable conditions these two gases could react to form $\text{N}_2(g)$ and $\text{CO}_2(g)$, which are components of unpolluted air.

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The carbon is oxidized because in the reactant, CO , is neutral however in the product CO_2 is also neutral even with the addition of another oxygen molecule. Therefore, the carbon must have lost a couple of electrons for the oxygen molecule to be attached.

$$b) K_p = \frac{[\text{N}_2][\text{CO}_2]^2}{[\text{NO}]^2[\text{CO}]^2}$$

$$c) \textcircled{i} \Delta G^\circ = \sum G^\circ_{\text{product}} - \sum G^\circ_{\text{reactant}}$$

$$= (-394.36 \text{ kJ}) - (2 \cdot 86.55 \text{ kJ} + 2 \cdot -137.15 \text{ kJ})$$

$$= -687.52 \text{ kJ}$$

$$\textcircled{ii} \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$\Delta S^\circ = \frac{-\Delta G^\circ - \Delta H^\circ}{T}$$

$$= \frac{-687.52 \text{ kJ} - (-746 \text{ kJ})}{298 \text{ K}}$$

$$= 0.196 \text{ kJ/K}$$

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$$d) \textcircled{i} \Delta G^\circ = -RT \ln K_p$$

$$= -(0.0821 \frac{\text{atm}}{\text{m}^2 \cdot \text{K}})(298) \cdot \ln(3.3 \times 10^{20})$$

$$= -24.5 \ln(3.3 \times 10^{20})$$

\textcircled{ii} To the left because since ΔG° is negative in the forward (right) direction, meaning it's not spontaneous, the left direction would mean spontaneous.